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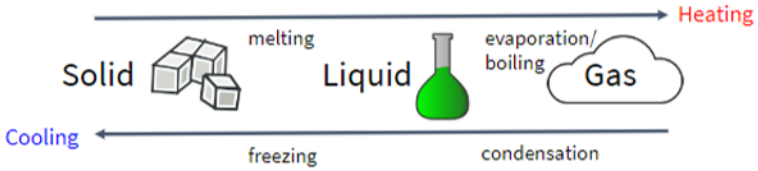
CHAPTER 4: STATES OF MATTER

1. Particles gain kinetic energy
2. Particles vibrate more vigorously
3. Forces of attraction between atoms/molecules weaken
4. Particles are further apart from each other



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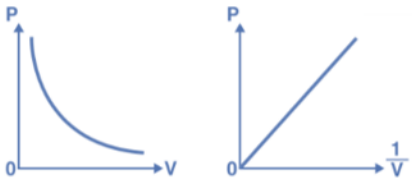
1. Particles lose kinetic energy
2. Particles vibrate less vigorously
3. Experience increasing forces of attraction
4. Particles are arranged more closely to each other

Assumptions of the Kinetic Theory of Gases:

- Gas molecules move rapidly and randomly
- All collisions between particles are elastic
  - No kinetic energy is lost due to collision
- Distance between gas molecules is greater than diameter of molecules
- No forces of attraction/ repulsion between molecules
- The temperature of the gas is related to the average kinetic energy of the molecules

Relationship between P, V and T:

- **Boyle's Law:**
  - At constant n and T, P is inversely proportional to V
  - $PV = nRT \rightarrow P_1V_1 = P_2V_2$



- **Gay-Lussac's Law:**
  - At constant n and V, P is directly proportional to T
  - $PV = nRT \rightarrow P_1/T_1 = P_2/T_2$

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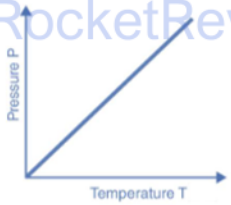

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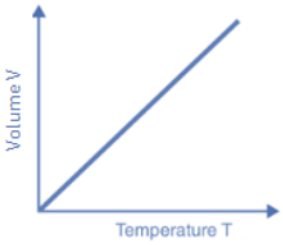
- **Charles' Law:**
  - At constant n and P, V is directly proportional to T
  - PV = nRT → $V_1/ T_1 = V_2/ T_2$
  - Standard temperature pressure: 22.4 dm^3 mol^{-1}
  - Room temperature pressure: 24.0 dm^3 mol^{-1}



- **Three Laws combine:** $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Ideal gas equation:

- PV = nRT
- n = PV/ RT

- M_r = g mol^{-1}
- M_r = m $(PV/ RT)^{-1}$
- M_r = mRT/ PV

	description	unit
P	pressure	kPa
V	volume	dm^3
n	number of moles of gas	mole
R	gas constant	8.31 J K^{-1} mol^{-1}
T	temperature	K

To prevent calculation errors, convert to the correct units first:
- Use kPa with dm^3
- Use Pa with m^3

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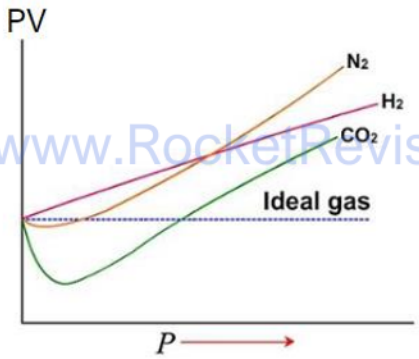
Ideal gas VS Real gas:

- **Ideal gas**
  - Volume of molecules is negligible
  - No forces of attraction/ repulsion between molecules
- **Real gas**
  - Volume of molecules is not negligible
  - Forces of attraction/ repulsion exist between molecules

Deviations from ideal gas behaviour:

- Deviations are observed at
  - **High pressure:** volume of molecules not negligible relative to container
  - **Low temperature:** intermolecular forces present, pulling molecules to each other



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Vapour Pressure:

- Pressure exerted by a vapour in equilibrium with a liquid, in a closed container
- Vapour pressure increases as:
  - Temperature increases
  - Gas particles have more kinetic energy
  - IMFs are overcome more easily
  - More gas molecules collide on the walls of the container

Partial pressure of a gas:

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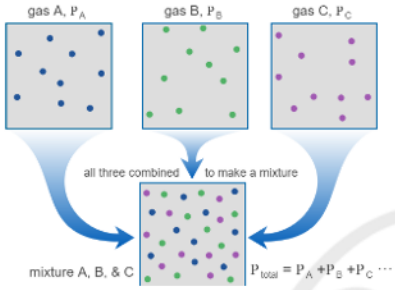
- Each gas in a mixture of gases exerts the same pressure that it would if it were alone and occupied the same volume
- $V_1P_1 + V_2P_2 = P(V_1 + V_2)$



- $Mole\ fraction = \frac{no.of\ moles\ of\ one\ gas}{total\ no.of\ moles\ in\ a\ mixture}$
- $Partial\ pressure = mole\ fraction \times total\ pressure$

Properties of liquids:

- Liquids have closer distances between molecules compared to gas
- Liquids have surface tension
  - Surface tension: cohesive forces between particles
  - Strong IMF lead to greater surface tension
  - Liquids tend to assume a geometry that minimizes surface area

Solids:

- Many ionic, metallic and covalent compounds are crystalline
- Lattice – a regularly repeating arrangement of ions/ atoms/ molecules in three dimensions

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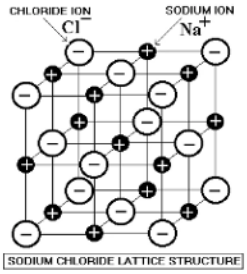
- **Giant ionic lattice:**
  - A 3-dimensional arrangement of alternative positive and negative ions
  - Ions held together by strong electrostatic forces
  - Properties:
    - Hard – strong electrostatic attraction between ions
    - Brittle
    - High melting and boiling points (higher charge density)
      Soluble in water: ion-dipole forces
    - Conducts electricity in molten or aqueous state (mobile ions)

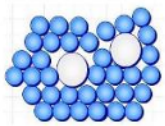


SODIUM CHLORIDE LATTICE STRUCTURE

- **Giant metallic lattice:**
  - A 3-dimensional arrangement of positively charged ions surrounded by a sea of delocalized electrons
  - Properties:
    - Malleable and ductile
      - Metallic bonds are easily broken and reformed
      - Delocalized electrons continue to hold the ions in the lattice
    - High tensile strength
      - Due to strong attractive forces between metal ions and delocalized electrons
  - Alloy:
    - Most metals are hard but some pure metals can be soft and malleable

• Layers of ions can slide over each other when a force is applied

    - Alloying can make a metal stronger by preventing the layers of ions from sliding over each other so easily



- **Giant covalent/ molecular lattice:**
  - Structures have a 3-dimensional network of covalent bonds throughout the whole structure
  - Properties are related to its structure
    - Diamond and SiO_2
      - High melting and boiling point – strong covalent bonds are broken
      - Hardness
      - Does not conduct electricity

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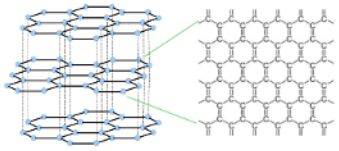
- E.g. Graphite, diamond, silicon (IV) oxide

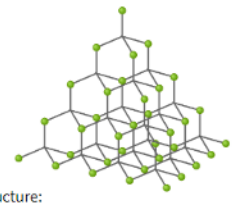
Graphite



Structure:
- Consists of sp^2 hybridized carbon atoms
- σ and π bonds present
- Carbon atoms arranged in hexagons in planar layers
- Layers of carbon atoms are held together by weak id-id forces

Properties:
- High melting and boiling point
  - strong covalent bonds are broken
- Soft
  - Weak id-id forces are broken
- Good conductor of electricity
  - Delocalised electrons present

Diamond



Structure:
- Consists of sp^3 hybridized carbon atoms
- σ bonds present
- Carbon atoms arranged tetrahedrally

Properties:
- High melting and boiling point
  - strong covalent bonds are broken
- Hard
- Poor conductor of electricity
- Good conductor of heat

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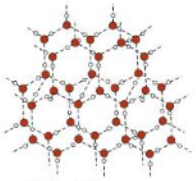
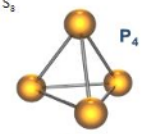
- **Simple molecular lattice**
  - A 3-dimensional arrangement of covalent molecules held together by weak intermolecular forces
  - The intermolecular force depends on the molecule:
    - Instantaneous dipole-induced dipole forces
    - Permanent dipole-permanent dipole forces
    - Hydrogen bonding
  - Properties
    - Lower melting/ boiling points

Sulfur, S_8



Water, H_2O



Phosphorus, P_4

Summary:

	Giant ionic	Giant covalent	Giant metallic	Simple molecular
Example	NaCl	Silicon dioxide	Copper	Iodine
Particles present	Na^+ and Cl^- ions	Si an O atoms	Cu^{2+} ions immersed in a sea of e^-	I_2 molecules
Forces keeping particles together	Ionic bonds	Covalent bonds	Metallic bonds	Temporary-induced dipoles
General physical state at room temperature	Solid	Solid	Solid	Solid
Melting and boiling points	High	High	High	Low

- Melting/ boiling point
  - Attractive forces between atoms/ ions/ molecules
  - Size of atoms/ molecules
  - Branching (organic chemistry)
- Electrical conductivity
  - Solid can conduct electrons in the presence of delocalized electrons (metals) and mobile ions (ionic compounds)

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- Solubility in polar or non-polar solvents
    - Depends if the solid is polar or non-polar
    - 'like dissolves like' – polar substances can dissolve in water

Exam answering tips:

When asked to explain the differences in melting / boiling points:

- Step 1: state the structure (giant ionic / giant metallic / giant covalent / simple molecular)
- Step 2: state the bonding (ionic / metallic / covalent / id-id / pd-pd / H bonding)
- Step 3: Compare and contrast the bonding
    - Strength of ionic bond ∝ charge density of cation (lattice energy : A2 syllabus)
    - Strength of metallic bond ∝ charge density of cation and no. of delocalized e^-
    - Strength of id-id forces ∝ no. of electrons per molecule
- Step 4: make the conclusion

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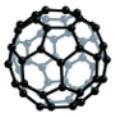
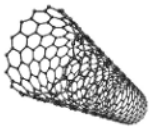
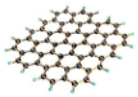
Allotropes of carbon:

- Diamond
- Graphite
- Graphene
- Carbon nanotubes
- Fullerene, C_{60}

Carbon nanoparticles:

	Buckminsterfullerene Fullerene (C_{60})	Carbon nanotubes	Graphene
Shape	Hollow sphere	Hollow tube	Isolated sheet of graphite
Hybridisation	sp^2	sp^2	sp^2
Electrical conductivity	Less conductive compared to CNTs and graphene	conductive	conductive
Special properties	Soft material Low sublimation point of 600 °C	High tensile strength along the long axis of the cylinder	Extremely strong for its mass

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